

Educational Analysis of the General, Organic, and Biochemical Laboratory

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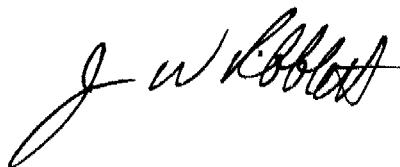
Educational Analysis of the General, Organic, and Biochemical Laboratory

An Honors Thesis (HONRS 499)

By

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Jason W. Ribblett

A handwritten signature in black ink, appearing to read 'J W Ribblett', with a large, sweeping initial 'J'.

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EDUCATIONAL ANALYSIS OF THE GENERAL, ORGANIC, AND BIOCHEMICAL LABORATORY

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Abstract

In order to provide every student with an equal and worthwhile experience in Chemistry, there must be many different modes of instruction especially for the laboratory experience. Experiments must not be entirely hands-on but also minds-on. Students cannot merely do the lab and leave without gaining any appreciation for the content. Experiments must be written in a way to make students think about what they are doing and learn from the experience. We have produced a laboratory manual that consists of twelve separate experiments written specifically for content retention. These experiments were written after careful consideration of behavioral theories, teaching techniques, and learning styles, such as Vygotsky's theories and Holistic Learning. For example, Vygotsky's theories provide the framework for the experiment, *Chemical Formula of an Oxide of Magnesium*, while the experiment *Le Chatelier's Principle and Buffers* was written using a Holistic Learning approach. Data from student surveys indicates that learning focused experiments provide better opportunities for content retention.

FOREWARD

Words cannot express my gratefulness for Dr. Jason Ribblett, for he has gone above and beyond his duty of research advisor. Over the course of our work together he has taught me a great deal about chemistry, teaching, and myself. Dr. Ribblett is a very humble person who puts his students first and leads by example. I can only hope that my future students will have the level of respect and gratitude for me that his former students have for him. I thank him again for his encouragement and support for chemistry education. Lastly, I thank him for always believing in me—as a chemist or as a teacher. I have found not only a mentor in Dr. Ribblett, but also a friend.

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I would like to thank my family for their never-ending and undying love and support. I would like to thank my fiancé Tyler Campbell for his support and enthusiasm for all that I do. His motivation has kept me believing in myself even in the worst of times. Last, I would like to thank my parents. There has never been a second of my life that my parents did not believe in me. I owe everything I do and achieve in life to them. My thesis and publication are both dedicated to them for all the love and support they have given me.

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Chemistry 101 is an introductory-level chemistry course designed for students pursuing degrees in areas of the health sciences. Most students who take Chemistry 101 are freshmen nursing students. The five credit hour course is a prerequisite for admission into the nursing program. Without a grade of C or better in Chemistry 101, students will not be admitted into the school of nursing. This rigorous course includes four credit hours of lecture and a two hour laboratory class each week. The students have a tremendous workload; for this course covers general, organic, and biochemistry in one semester. A student majoring in chemistry would have four semesters to cover the same material covered in this one class. Chemistry majors take two semesters of general chemistry as well as separate classes of organic chemistry and biochemistry.

With such a heavy workload, it is easy for students to fall behind, lose hope, and give up in this class. For some, this is the first chemistry class they have ever taken. For others, this is the first chemistry lab they have ever entered. For many of the students, the mathematical demand makes the class seem impossible. These students want to be nurses. They want to be nurses because they have the necessary characteristics of a nurse: compassion, patience, nurturing. Some never make it past the first semester of college in the nursing program because this new demanding subject beats them. The rigorous nature of this course has prevented many capable and great potential nurses from entering the field. The main goal of my thesis project is to prevent this from happening to future nursing students.

The problem that I have been working with is, "How can the Chemistry 101 professors help students learn and, thus, succeed in this class?" One area that needed serious attention was the lab. The lab portion of any science class serves many purposes. One major advantage to a lab is that it provides hands-on and visual opportunities for students to experience the content they are studying in lecture. Labs are also smaller in number than lectures, so there is a smaller ratio of student to instructor. This allows students to receive more one on one attention from a professor or teaching assistant. Laboratory experiments are designed as a supplement to lecture. Students

are able to observe chemical phenomena and ask their lab partners or instructors questions. In theory, labs are wonderful teaching tools.

Labs can easily lose their effectiveness, which is what happened to the Chemistry 101 lab. The lab manual chosen for this lab was one which had accompanied the textbook and was written by a collection of professors from around the country. It is a good manual, but it was not written specifically for Ball State University's Chemistry 101 course. Some experiments were not written the way the professors preferred. In fact, they had to rewrite some experimental procedures before the students could perform them because certain parts did not pertain to the lecture material. Many of the experiments were written in a way so that the students did not have to think critically. They were written too much like a recipe, where the students would perform the experiment step by step without ever thinking about what they were doing. The use of this manual separated the lab and the lecture. From the student's perspective, some experiments seemed to have no connection with the lecture material. Students were not using the lab as a tool for the class; instead they saw it more as a burden. At times, students would leave the lab and an hour later not know what they had done. Students saw little connection between the lab and the lecture; they were not being challenged; and they were not using the lab to improve their comprehension in the class. The lab was not being used as it should have been.

To help the students learn better, the professors of this course decided to write their own lab manual. Having an in house lab manual would tailor the lab portion of Chemistry 101 to the lecture. All the material in the manual would be pertinent to the course. Everything the professors wanted the students to know and experience would be found in the manual. In addition to the connection to lecture, this lab manual would contain experiments that challenged the students to think critically. Another goal of this new manual was to provide a study tool for students.

I belonged to a team of student researchers who assisted two professors in the production of this lab manual. This project began two years before I entered the lab, but only one experiment had been completed in its entirety. Our main goal was to publish a manual with as many as twelve experiments for use in the spring semester of 2005. We wanted to design these labs in a way that the students had to think critically and connect the lab to the lecture. Again, we wanted these labs to serve as a study tool, for the lab should complement the lecture.

I worked extensively on two experiments, "The Chemical Formula of an Oxide of Magnesium" and "Le Chatelier's Principles and Buffers" (Appendices A & B). The first question I asked myself when I began researching was, "How do I make these experiments more meaningful than the similar experiments in the old manual?" My goals reflected that question. My first goal was to connect the lab not only to the lecture, but to real life. I wanted these experiments to have relevance in the students' lives. My second goal was to structure these experiments to give the students an organized way to study for the class.

Writing a laboratory experiment is the easy part. There are five sections: **Background**, **Procedure**, **Data Sheet(s)**, **Pre-Lab Questions**, and **Post-Lab Questions**. Background information is written in the **Background**. This information should be read by the student before entering the lab. It contains information similar to the lecture notes and the textbook. The background is lab specific. It contains examples from the upcoming lab and explains the relevance of the lab. From the background, the lecture, and the textbook, the students should be able to answer the **Pre-Lab Questions**. These questions cover the topic for the lab. It is important to have pre-lab questions because this ensures that the students have read the background and have a general understanding of the content before entering lab. Pre-lab questions can be written to get the students to focus on a specific concept. These questions should be completed by the students before entering the lab. Next, the students follow the **Procedure**. This explains the steps of the lab, points out safety cautions, and gives the structure

of any calculations. Students follow the steps of the procedure and record their observations on the **Data Sheet**. After the lab, students complete the **Post-Lab Questions**. These questions are generally application and analysis questions drawing from data collected in the lab. Again, writing a laboratory experiment is easy. Writing a lab that is meaningful to the students is a challenge. It was my challenge.

In order to make these five parts of the lab more meaningful, I researched different learning styles and teaching techniques. It was important to understand how students learn. As I began my project, I focused heavily on the research of Vygotsky and the theory of holistic learning.

Vygotsky, a Soviet psychologist, believed that learning was a lifelong process that depended on social conditions and connections. Vygotsky believed that learning must be achieved with peer or adult guidance because the zone of proximal development cannot be overcome independently. The zone of proximal development describes the distance between what a student already knows and what a student has the potential for knowing. To reach this unknown knowledge, a student needs the guidance of a teacher or peer. He called this guidance *scaffolding*. Teachers must provide scaffolding in order for students to successfully cross the zone of proximal development. Once a student has acquired the new knowledge, he or she has made a discovery. Allowing students to make their own discoveries is a great teaching tool. Vygotsky believed that instead of telling students concepts and answers, students should be allowed to discover for themselves, because the content will then be internalized if the students make their own discoveries. The content will have much more meaning to the student because he or she made an independent discovery. Vygotsky believed that students could not make these discoveries without the proper scaffolding. Teachers must set up proper structure in order for students to be capable of making a personal discovery.

A lab is an ideal situation for personal discoveries. However, it is not wise to give students a bench full of chemicals and say, "Here, discover." This is where scaffolding is very important.

Students need the background structure so they can make wise and safe choices during lab. I used the concept of scaffolding when designing the lab, “The Chemical Formula of an Oxide of Magnesium.”

The Magnesium Oxide lab is the second experiment that the students perform. This lab is used to explain the concept of stoichiometry. Stoichiometry is a very challenging and foreign area of chemistry for many new students. It takes practice and determination to successfully master this math-driven part of chemistry. Many students do not grasp the concepts fully at the beginning, which creates much stress, causing some students to give up. This lab gives the students practice performing stoichiometric calculations. Students burn strips of magnesium in a porcelain crucible in the presence of air, which contains oxygen. Ultimately, the magnesium and oxygen react to form a compound, magnesium oxide (MgO). It is not difficult for students to understand what compound results from this reaction (magnesium oxide). The difficult part is predicting the correct chemical formula of the compound (MgO). Students must use data from their lab to correctly label their product as MgO , not Mg_2O , nor Mg_3O_2 . They must be able to calculate the 1:1 ratio between the magnesium atoms and the oxygen atoms in this compound. These calculations form the basis of many stoichiometry problems.

Since this topic of stoichiometry is still new to the students, it is essential that this lab be written with the zone of proximal development in mind. Students will know the basic theory of stoichiometry when entering the lab, but they will not be extremely confident with their skills and understanding of the subject matter. Using the lab as a means to cross the zone of proximal development, these students should be able to leave the lab with a better understanding of stoichiometry. They will have discovered how to use data from a lab to calculate the formula of a compound which they created. This is an experience the students can recall when they do other practice problems at home or in class. They can review the lab calculations while remembering the procedure they used. This is a personal discovery.

This personal discovery cannot be achieved without the necessary scaffolding. In the Magnesium Oxide lab, I included scaffolding in the **Background**, **Pre-Lab Questions**, and **Data Sheet**, with each component having less and less scaffolding. In the end, the students achieve their own discovery. In the **Background**, there is a description of an experiment very similar to the one they will perform. The example describes in detail each step of an experiment in which a student is determining the formula of a new compound. This example also walks students through the necessary calculations. The calculations are shown using an organized routine. Students can read through this structured example and understand the final answer. After reading the **Background** and before coming to lab, students complete a pre-lab question that is quite similar to the example in the **Background**. This question presents a scenario very similar to the example in the background and the procedure they will be performing later. The students are asked to determine the formula of the compound in the pre-lab question. The question is organized in a way that the students see the order in which they should perform the calculations, but they are not told how to calculate the answer. They must refer to the more structured background for this information. The students will determine the formula of the compound in the pre-lab question and make a discovery, but it is not completely personal yet. It is not until they follow the **Procedure** and answer the **Post-Lab Questions** that the students complete the discovery process.

After students have completed the **Procedure** and filled in their **Data Sheet**, they are asked to determine the formula of the compound using their data. It is in this step that the students are crossing the zone of proximal development. They are making a personal discovery of the formula of their compound using the scaffolding we set for them as support. Students use the structure set up for them in the **Background** and **Pre-Lab Questions** to assist them in the determination of their compound's formula. This is now a personal experience that they can relate to as they continue their learning of stoichiometry.

By using a scaffolding approach, teachers are setting up their students for success. Again, this is a subject that does not always ensure student success. Many students lack the organizational and mathematical skills that stoichiometry demands. Using scaffolding is a way of helping them understand stoichiometry while improving their organizational skills. Once students can organize their calculations, the math is less challenging. Once the math is no longer the challenge, the problem is less difficult to master. As students conquer more and more of these problems, their confidence increases, and stoichiometry suddenly is not so difficult.

Scaffolding is a vital part of instruction. It is also important to make connections whenever possible. Vygotsky would agree that connections are essential, for his theories explained that *learning* was a connection between individuals. It is important to connect units in chemistry to each other. Students must be made aware of how different parts of chemistry relate. In addition, students must be shown connections of chemistry with other subjects. Chemistry should not be isolated. It should integrate other sciences, math, English, history, and as many other content areas as possible. It should have connections with every day life as well. If students can see how chemistry plays a role in their personal lives, it will become more of a personal subject of interest. Connections are important. This is the basis of holistic learning.

Holistic learning uses connections to engage students. It is organized in a method called “whole-part-whole.” Teaching holistically begins with the “whole.” During this segment, students use their prior knowledge, however much or little that may be. Tapping prior knowledge is a way of connecting the lesson to the students’ experiences. It can be as simple as connecting a soda a student drinks to acids and bases or as complex as connecting the respiratory cycle to the acidity of a person’s blood. Making these connections early on is the “hook” for the students. The “whole” part of holistic learning also includes a general overview of the lesson or unit. In the case of an experiment, this is included in the **Background** and **Pre-Lab Questions**. The **Background** gives a general overview of the topic of the experiment, and the **Pre-Lab**

Questions are often general in nature, requiring students to use their prior knowledge of the subject.

After the general overview, teachers can start the next segment of holistic instruction, the “part.” During this segment, the content is broken into small, detailed parts. Each part is covered specifically—always making connections to other parts and to the whole. For example, in an acid and base unit, there could be several different parts discussed in length: theories, acid and base strength, compound naming, reactions, pH, titration, acid rain, etc. It is important that none of these lessons be isolated. They should always be connected to the previous lessons and to future lessons. In an experiment, the **Procedure** serves as the “part” component. Students are performing an experiment one step at a time and observing all the details of the different parts individually. It is impossible to truly separate and isolate all of these parts. The students should be observing the connections, but often they only see isolated parts. It was my challenge to write experiments that made connections clear to the students.

Finally, it is crucial to remember the last segment of holistic learning, the final “whole.” During this segment, a final connection between all the parts is made. Students should be able to summarize the unit. They should be able to analyze data and think critically about the unit. They should be able to apply the content and identify connections to other classes or life situations. This is a time of deep thinking that is vital for the retention of information. In an experiment, this occurs during the **Post-Lab Questions**. The **Post-Lab Questions** are used for application and analysis purposes. Students must apply what they observed in the lab to answer the questions.

When writing the lab for Le Chatelier’s Principle, I relied heavily on principles of holistic learning. Le Chatelier’s Principle explains chemical equilibria, another topic that is challenging to beginning chemists. Le Chatelier’s Principle explains how systems in equilibrium will react if a stress is applied to the equilibrium. It is one thing to read about Le Chatelier’s Principle. It is

another to actually observe this principle in action. It is very important to let students observe the visual triggers that help explain the principle.

When writing this lab, I focused on predictions, observations, and conclusions. These three elements are important parts of holistic learning. Students should be using their prior knowledge to predict (whole) what they may eventually observe. After making the observation (part), the students must be able to make connections by summarizing (whole) what they saw and relate that to what they predicted. I incorporated a table into the Le Chatelier's Principle lab that has students predict, observe, and summarize.

For one of the **Pre-Lab Questions**, students are to complete two columns of this table, the Shift column and the Prediction column. Both of these columns rely on the prior knowledge of the student. Students must use their knowledge from the **Background**, lecture notes, and the textbook to predict how a system in equilibrium will shift when a certain stress is applied. Then they must write a prediction of what they expect to observe. Will the system change color? Will a solid appear in the system? Will a gas be produced? Will heat be given off? The prediction column is very important, not only because it requires students to use their prior knowledge, but also because it includes the nature of science. This exercise is designed to show students that in science, sometimes predictions are wrong. Nonetheless, making predictions is very important. These two columns serve as the “whole” part of the lab instruction.

As students perform the experiment, they apply different stressors to several systems in equilibrium and observe what happens. For example, applying the stress may cause the solution to change color. During this “part” segment, students are to complete the Observation column of the table. Students are to write details of what they observed as the stress was applied. After this specific and detailed portion of the lab, students are to complete the Summary column as a post-lab question. In this column, students compare what they observed with what they predicted. This is where students must be reflective and analytical. In some cases, their observations and

predictions will agree. In other cases, they could be very different. The students must provide insight as to why the observation was different from the prediction. They must explain what they have learned. They must connect the part to the whole during this last “whole” part of instruction. This also reinforces that it is acceptable to be wrong in the beginning. What is important is that something was learned.

How effective *were* we in getting students to learn? We surveyed the students from the Fall 2004 course who used the old manual as well as the students from Spring 2005 who used the new manual (Appendix D). These surveys, administered during the final exam period, asked questions pertaining to our three main objectives: lab being enjoyable, lab being a learning experience, and lab being a study tool. Data analysis shows us that we had success meeting two of the three objectives: helping students enjoy lab and helping students learn.

Students were asked the question, “Did you find the laboratory experiments to be enjoyable?” with possible answers of “no,” “sometimes,” “most of the time,” and “always.” There was a 5% increase in those who answered positively (most of the time or always) from fall (52%) to spring (57%). We attribute this increase to our changes, especially those we made to our procedures and pre-lab questions. Another question on the survey was, “Were the procedures easy to read and understand?” with response choices of “never,” “sometimes,” “most of the time,” and “always.” In the fall, 13% of students replied “never.” In the spring, no student replied “never.” In addition, 89% of spring students replied “most of the time,” and “always,” a 19% increase from the fall semester. Having clearly explained procedures relieves the stress of being confused. Students are then able to focus on what they are doing instead of being confused as to what to do. We also asked students, “Did the pre-lab questions help you prepare for the experiment?” with four answers: “never,” “sometimes,” “most of the time,” and “always.” There was an 8% increase of students who answered positively (most of the time or always) from fall (33%) to spring (41%). Pre-lab questions are vital for the holistic approach of laboratory

instruction. Although there was an increase in students who found them useful, we will continue to modify these questions in order to further improve the manual.

We asked the students which experiment was the most enjoyable of the semester. The results from the fall were interesting. Out of the top four answers, three of those experiments were experiments we had written and given to the students as handouts. The fall labs provided a way for us to test run a few of our new experiments before publishing them. Of twelve possible experiments, 50% of students answered that the Saponification: The Process of Making Soap, an experiment we designed, was their favorite. The next favorite experiment was also one we designed; the Esters experiment received 20% of the votes. Tied for third at 12% each were the Enzyme experiment which we designed and the Fruit Juice Titration experiment which was part of the former laboratory manual. These overwhelming percentages of students who greatly enjoyed the three experiments we handed out caused us to raise an important question. If we had not handed out our three test experiments, would the percent of students who answered that they found the laboratory experiments to be enjoyable “most of the time” or “always” be lower? The percentage of students who greatly enjoyed our experiments leads us to believe that there is a link between the handout experiments and the enjoyable quality of the lab.

In the spring, the Saponification: The Process of Making Soap experiment was again the favorite with 63% of students claiming it was the most enjoyable. The Esters and Enzyme experiment were the second and third favorites with percentages of 25% and 8% respectively. It is clear that these three experiments were enjoyable to the students.

It is not enough that the students enjoyed lab, but also that they learned in the process. Our data leads us to believe that comprehension was higher in the spring than in the fall. We asked students if they had ever taken chemistry before prior to this class. In the fall, 7.5% of students had never taken chemistry while in the spring, 15% of students had never taken chemistry. Even with the higher percentage of new chemistry students, the spring students show higher signs of

comprehension. Only 1% of the students in the spring section of CHEM 101 failed the laboratory portion of the class; whereas, in the fall 5% of the students failed the laboratory portion. The average class grade has been higher throughout the entire spring semester than in years past. Although this cannot be attributed entirely to the laboratory portion of the class, the lab cannot be dismissed entirely as part of the reason either.

Our survey contained a series of questions pertaining to the experiments. These were content knowledge questions that students should be able to answer correctly based on concepts presented in the lab. Four of six questions showed an increase in the percent of students who answered correctly from fall to spring, while one question had nearly equal percentages in the fall and spring. There is one question in particular with increasing percentages that we accredit to our laboratory manual. The correct answer to the question "Which of the following glassware is most accurate at measuring liquid volumes?" is either a buret or pipette. Other choices are a beaker or a graduated cylinder. In the fall, 56% of students answered correctly, while a shocking 7% of students answered beaker, a very inaccurate means of measuring volume. In the spring, 63% of students answered buret while only 3% of students answered beaker. In the new laboratory manual, glassware accuracy is tested in the Density Lab.

Our main objective was using the lab as a study tool. We asked students, "Did you use the lab experiences or the lab write-ups to help you prepare for the exams?" Students had four choices of answers, "never," "sometimes," "most of the time," and "always." Although we saw a 3% increase from fall to spring in those who use the lab always to study, we saw an 8% increase in those who never do. In the spring semester, 54% of students answered they never used the lab as a study tool. Although we did not see the results for this objective that we were hoping to see, we now have a second area of focus, in addition to modifying pre-lab questions. It is important that students enjoy the lab and learn from the lab, but it is also important that they use the lab to help them study. Our main goal for the next edition of the lab manual is to improve this area by

finding a way to further connect the lab to the lecture so students see the lab as a valuable study tool.

This research project served as a major renovation for the CHEM 101 course. The results from our surveys (Appendix E) keep us optimistic that we have made a positive impact on the student learning. Although we have made many changes, there are still many more changes which need to be made before we see a definite improvement in the lab. I feel that the introduction of holistic lab instruction, as well as the implementation of student surveys, will allow future researchers to adapt this course to accommodate students as the need arises.

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Appendix A

The Chemical Formula of an Oxide of Magnesium

The Chemical Formula of an Oxide of Magnesium

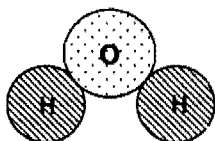
Purpose of the Experiment:

Determine the chemical formula of an oxide of magnesium.

Background

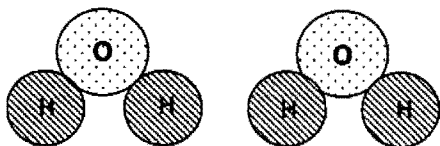
When two or more different atoms combine to form a pure substance, it is called a compound. There are two types of compounds, molecules and ionic compounds. You will learn more about each of these as the semester progresses. The composition of either type of compound is symbolized by its chemical formula. A chemical formula includes the atomic symbol of each element present in the compound followed by a subscript specifying the number of atoms present. If a subscript is absent, only one atom of that element is present in the compound.

A water molecule contains two atoms of hydrogen and one atom of oxygen. Thus, water has the chemical formula, H_2O .



One water molecule contains two hydrogen atoms and one oxygen atom.

The subscripts act as an atom ratio. In each water molecule, there is a 2:1 ratio of hydrogen atoms to oxygen atoms. If we were to increase the number of water molecules, the ratio would still remain at 2:1, even though the total number of atoms changes. For example, two water molecules would consist of four hydrogen atoms and two oxygen atoms, ten water molecules would consist of twenty hydrogen atoms and ten oxygen atoms, 1000 water molecules would consist of 2000 hydrogen atoms and 1000 oxygen atoms, and 6.022×10^{23} water molecules would consist of 12.044×10^{23} hydrogen atoms and 6.022×10^{23} oxygen atoms.



Two water molecules contain four hydrogen atoms and two oxygen atoms.

The value 6.022×10^{23} is called Avogadro's number and is often referred to as a **mole**. Just as there are twelve doughnuts in one dozen of doughnuts, there are 6.022×10^{23} atoms in one mole of atoms. Thus, we can rewrite " 6.022×10^{23} water molecules would consist of 12.044×10^{23} hydrogen atoms and 6.022×10^{23} oxygen atoms" as "*one mole* of water molecules would consist of *two moles* of hydrogen atoms and *one mole* of oxygen atoms". Because a water molecule has an atom ratio of 2:1 (two hydrogen atoms for every one oxygen atom), it also has a mole ratio of 2:1 (two moles of hydrogen for every one mole of oxygen). The subscripts in a chemical formula not only give the atom ratio of the atoms in the compound, they also give the **mole ratio** of atoms in the compound. Chemical formulas represent the mole ratios of the atoms in a compound.

The mole is very useful because it also relates the number of atoms in a sample (a number too large for you to count in your lifetime) to the mass of those atoms in grams (it only takes a few seconds to weigh a sample on a balance). The mole can be defined as the number of atoms present in a 12.01 g sample of carbon. A 12.01 g sample of carbon contains 6.022×10^{23} carbon atoms, or one mole of atoms. Thus, one mole of carbon weighs 12.01 g. The mass of one mole of a substance is called a **molar mass**. The molar mass of most elements is equal to its atomic mass expressed in grams. The atomic mass of calcium is 40.08 amu; thus, one calcium atom weighs 40.08 amu. By definition, the molar mass of calcium is 40.08 g/mol. This means that a 40.08 g sample of calcium contains one mole, or 6.022×10^{23} atoms.

The molar mass can be used to interconvert between the number of moles and the mass of an element. To find the mass of a certain number of moles of an element, simply multiply the number of moles by the element's molar mass. Because 1.0 mole of calcium has a mass of 40.08 g, 1.5 moles of calcium would have a mass of 60.12 g ($40.08 \text{ g/mol} \times 1.5 \text{ mol}$). To find the number of moles of an element from its mass, simply divide the mass by the molar mass. Because a 40.08 g

sample of calcium consists of 1.0 mole of calcium, 80.16 g of calcium would contain 2.0 moles (80.16 g ÷ 40.08 g/mol).

The molar mass of a compound can be determined from its chemical formula because the chemical formula of a compound represents the mole ratio of atoms in that compound. One mole of water molecules consists of two moles of hydrogen atoms and one mole of oxygen atoms. The mass of two moles of hydrogen is 2.02 g, and the mass of one mole of oxygen is 16.00 g. Thus, the mass of one mole of water is 18.02 g, making the molar mass of water 18.02 g/mol.

Chemical formulas do not, however, represent the mass ratio of atoms in a compound. The mass ratio of hydrogen to oxygen is 1:8; whereas, the mole ratio is 2:1. Because a chemical formula represents a mole ratio, it is necessary to convert any masses determined in the lab to moles. Then, a mole ratio can be determined. Using the mole ratio, a chemical formula can be assigned.

Take the following as an example. A student heated 0.70 g of nitrogen in an open container and produced a compound of nitrogen and oxygen. The compound of nitrogen and oxygen had a mass of 2.3 g. From this information, the chemical formula can be determined using the following steps.

1. To begin, the mass of each element in the compound must be determined. Based on the information given, there are 0.70 g of nitrogen present. Because only one other element is present in the compound, the mass of oxygen can be determined by subtracting the mass of nitrogen from the mass of the compound.

$$2.3 \text{ grams} - 0.7 \text{ grams} = 1.60 \text{ grams oxygen}$$

2. To determine the number of moles of nitrogen and oxygen in the compound, divide the mass of each element by their respective molar mass.

$$\text{moles of N} = 0.70 \text{ g N} \left(\frac{1 \text{ mol N}}{14.01 \text{ g N}} \right) = 0.050 \text{ mol N}$$

$$\text{moles of O} = 1.60 \text{ g O} \left(\frac{1 \text{ mol O}}{16.00 \text{ g O}} \right) = 0.100 \text{ g O}$$

3. To calculate the mole ratio, the number of moles of each element must be divided by the smaller number of moles. In this example, the 0.050 moles of nitrogen is used because it is the smaller number of moles.

$$\text{N ratio} = \frac{\# \text{ mol N}}{\# \text{ mol N}} = \frac{0.050 \text{ mol N}}{0.050 \text{ mol N}} = 1.0$$

$$\text{O ratio} = \frac{\# \text{ mol O}}{\# \text{ mol N}} = \frac{0.100 \text{ mol O}}{0.050 \text{ mol N}} = 2.0$$

There are two moles of oxygen for every one mole of nitrogen present in this compound. Therefore, the chemical formula would be NO_2 .

Sometimes mole ratios do not contain whole numbers. Many compounds have ratios such as 1:1.5. In these cases, ratios must be multiplied by a common factor in order to get two whole numbered subscripts. For the ratio 1:1.5, if both numbers are multiplied by 2, the ratio becomes 2:3. Table 1 lists some common ratios and their respective whole number ratios.

Ratio	Common Factor	Whole Number Ratios
1:1.67	3	3:5
1:1.50	2	2:3
1:1.33	3	3:4
1:1.25	4	4:5

Table 1

In this lab, you will heat a magnesium sample to form a compound of magnesium and oxygen. Using calculations like those outlined above, you will determine the chemical formula of the oxide of magnesium.

Procedure

- A. Attach an iron ring to a ring stand, allowing enough room underneath for a Bunsen burner. Place a clay triangle on the ring.

Note: Always use a clean, dry crucible. If they are not clean, masses will not be accurate, and results will contain error. Always use crucible tongs to handle the crucible and cover to prevent the transfer of fingerprints. Never touch a hot crucible.

- B. Place a porcelain crucible at a slight angle on the triangle. Place the crucible cover on the crucible so that it is slightly ajar. See **Figure 1**.



Figure 1

- C. Turn on the gas and light the Bunsen burner. Adjust burner to produce a very hot flame 1-2 inches in height.
- D. Heat the crucible gently for 5 minutes by moving the flame back and forth across the bottom of the crucible. After 5 minutes, turn off the gas to extinguish the flame.
- E. Remove the cover from the crucible. Set it on wire gauze to cool. Next, remove the crucible, and allow it to cool on the wire gauze.

Note: Never weigh a hot object, for the mass will not be accurate. To determine if an object is cool enough, hold your hand about 2 cm from the object. If no heat is felt on your hand, it is cool enough to weigh.

- F. When the crucible is cool, use crucible tongs to transfer the crucible and cover to the balance. Determine the mass of both together. Record this mass on line **F** of Data Sheet I.

- G. Obtain 3 strips of magnesium ribbon. Loosely roll the magnesium ribbon into a ball, and place it in the crucible.

- H. Weigh the crucible, magnesium, and cover together on the balance. Record this mass on line **H** of Data Sheet I.

- I. Place the crucible with magnesium at a slight angle on the triangle with the cover slightly ajar, as in **Figure 1**.

- J. Turn on the gas and light the Bunsen burner.

Caution: Never look directly into the white flame of burning magnesium, for Mg emits UV light when burned. Eye damage could result.

Note: Completely cover the crucible at the first sign of white smoke. This white smoke is the product, the oxide of magnesium. Allowing it to escape will cause error in your results.

- K. Gently heat the crucible by slowly moving the flame back and forth across the bottom of the crucible. If white smoke is produced, remove the flame; and cover the crucible **completely**.
- L. After about 15-20 seconds, lift the cover slightly, and check to see if white smoke is still present. If smoke is present, repeat steps J through L until you no longer observe smoke. You may have to repeat these steps many times. Proceed to step M when no smoke appears.
- M. Place the cover so that it is slightly ajar, as in Figure 1. Light the Bunsen burner, and place it under the crucible so that the flame contacts the lower edge of the crucible. Heat the crucible with a strong flame until the bottom of the crucible is red. Heat for 10 minutes.
- N. Turn off the gas to extinguish the flame.
- O. Remove the cover, invert it, and allow it to cool on the wire gauze. Let the crucible cool on the triangle until no heat is felt when you hold your hand about 2 cm away.

Caution: Before adding the water in step Q, let the crucible cool to room temperature. Adding water to a warm crucible will cause it to shatter and could result in serious injury.

- P. Add 10 drops of deionized water to the crucible.
- Q. Replace the cover on the crucible so that it is slightly ajar.
- R. Turn on the gas, and light the Bunsen burner.
- S. Heat the crucible for 5 minutes by slowly moving the flame back and forth across the bottom of the crucible.
- T. Place the lighted Bunsen burner under the crucible so that the flame is in contact with the bottom of the crucible. Heat the crucible strongly for 5 minutes.
- U. Turn off the gas to extinguish flame.
- V. Remove the cover, invert it, and allow it to cool on the wire gauze. Also, remove the crucible and place it on the wire gauze to cool.
- W. When crucible and cover are cool, find the mass of the crucible, its contents, and the cover. Record this mass on line W of Data Sheet I.
- X. Discard your product according to your instructor's directions. Wash both the crucible and cover with tap water; rinse with deionized water.

Name _____

Date _____

Post-Lab Questions

1. Why should you wait until the crucible is completely cool before weighing?

3. Use the following steps to determine your percent error for the molar mass of magnesium oxide.

a. Multiply the mole ratio values of Mg and O by their respective molar masses.

$$\text{_____ Mg ratio} \bullet 24.31 \text{ g/mol} = \text{_____ g/mol}$$

$$\text{_____ O ratio} \bullet 16.00 \text{ g/mol} = \text{_____ g/mol}$$

2. How would your Mg to O ratio have been affected if:

Add these two amounts together in order to calculate the molar mass.

a. the Mg was not burned completely?

_____ g/mol

b. The accepted molar mass of magnesium oxide is 40.31 g/mol. Determine your percent error using the following equation:

b. the cover was lifted and smoke, the oxide of magnesium, escaped from the crucible?

$$\frac{(\text{accepted mass} - \text{calculated mass})}{(\text{accepted mass})} \bullet 100\%$$

_____ %

Name _____ Date _____

Data Sheet

F. mass of crucible and cover _____ grams

H. mass of crucible, cover, and magnesium _____ grams

W. mass of crucible, cover, and magnesium oxide _____ grams

Calculations:

1. Mass of magnesium (**H - F**) _____ grams

2. Mass of oxygen (**W - H**) _____ grams

3. Moles of magnesium (molar mass of Mg is 24.31 g/mol) _____ moles

4. Moles of O (molar mass of O is 16.00 g/mol) _____ moles

5. Mole ratio of Mg (number of moles of Mg to the smaller of the two amounts of moles) _____

6. Mole ratio of O (number of moles of O to the smaller of the two amounts of moles) _____

7. Chemical formula for the oxide of magnesium _____

Pre-Lab Questions

1. Which contains more atoms: 8.00 grams of beryllium or 80.0 grams of silver? Show calculations to support your answer.
- c. Determine the number of moles of O in the compound.

_____ moles

- d. Find the mole ratio of Fe in the compound.

2. When objects made of iron, Fe, are left in the open air, they will rust. Rust, an oxide of iron is formed when the iron reacts with oxygen in the air. When 3.50 g of iron was left in the air, eventually 5.01 g of rust formed. The molar masses of iron and oxygen are 55.85 g/mole and 16.00 g/mole respectively. Find the chemical formula of this oxide of iron using the following steps:

- a. Determine the mass of O in the compound.

_____ grams

- b. Determine the number of moles of Fe in the compound.

_____ moles

- e. Find the mole ratio of O in the compound.

- f. Write the chemical formula for the rust formed, referring to Table 1 if necessary.

Appendix B

Le Chatelier's Principle and Buffers

Le Chatelier's Principle and Buffers

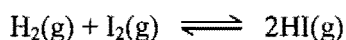
Purpose of the Experiment:

Make predictions based on the concepts of Le Chatelier's principle and the Henderson-Hasselbalch equation. Observe changes in equilibrium as concentrations of reagents or products are changed. Monitor changes in pH upon addition of acid or base to water and buffer solutions.

Background

Le Chatelier's Principle

Many reactions are **reversible**. A reversible reaction is one that can go in a forward and backward direction, meaning that sometimes the products of a reaction can react to form the reagents. Reversible reactions are written with a double arrow:



If we mix hydrogen gas and iodine vapor in a flask at 200 °C, only the forward reaction can occur because no product, HI, has been created yet. As the reaction continues, there will be decreasing amounts of hydrogen gas and iodine vapor, causing the rate of the forward reaction to decrease. As the product, HI, is produced, the reverse reaction will begin to occur, with 2HI being the reagents, and hydrogen gas and iodine vapor being the products. The rate of the reverse reaction will increase as the amount of HI present increases. The rate of the forward reaction will continue to decrease, and the rate of the reverse reaction will continue to increase until the two rates become equal. When this occurs, the reaction is said to be in a state of **equilibrium**.

Do not confuse this concept of equal rates with equal concentrations. Although a reaction in equilibrium has forward reaction and reverse reaction rates that are equal, it does not mean that the concentrations of the products are equal to the concentrations of the reagents. A reaction at equilibrium could have very high concentrations of reagents, and low concentrations of products, yet still be at equilibrium. Once a reaction has reached equilibrium, the concentrations of the reagents and the products will not change, even though they may be different. If the concentration of a certain reagent at equilibrium is 0.4 M, it will remain at 0.4 M as long as the reaction remains at equilibrium. The reaction could stay at equilibrium for two seconds or two hundred years; but as long as it is at equilibrium, the concentration of that reagent will be 0.4 M.

Chemical equilibria can be disturbed by outside forces, or stressors. If a stress is applied to a system already at equilibrium, the position of equilibrium will **shift** in a direction that will relieve the stress. This concept is known as **Le Chatelier's principle**. Changes in concentrations of reagents or products, changes in temperature, or changes in pressure are all examples of stresses that could disturb equilibria. In this lab, Le Chatelier's principle will be examined by changing the concentrations of the reagents or products.

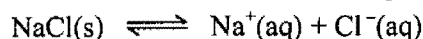
Part 3 of the Le Chatelier's Principle section involves the use of color changes to observe shifts in chemical equilibria. A colorimeter will be used to confirm the color changes. A colorimeter uses spectroscopy to measure the absorbance of solutions. When a solution is placed in the colorimeter, light passes through the solution, and strikes a detector. The detector measures how much light is absorbed by the solution. According to Beer's Law,

$$A = k \cdot C$$

absorbance (A) has a direct relationship to concentration (C), meaning that the amount of light a solution absorbs is directly proportional to the amount of solute contained in the solution. Thus, changes in the concentrations of the reagents or the products can be monitored by measuring fluctuations in absorbance.

Part 1: Equilibrium of a Saturated NaCl Solution

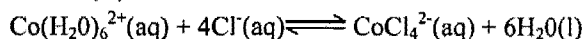
A saturated saltwater solution has the equilibrium:



Adding additional chloride ions from HCl would disturb the equilibrium by increasing the concentration of a product. To relieve the stress, the position of equilibrium will shift to the left (toward reagents). Thus, an increase in solid NaCl should be observed.

Part 2: Equilibrium of $\text{Co}(\text{H}_2\text{O})_6^{2+}$ and CoCl_4^{2-}

When dissolved in water, a cobalt (II) ion will bond to six water molecules to form the complex ion, $\text{Co}(\text{H}_2\text{O})_6^{2+}$. The resulting solution has a pink color. Addition of chloride ions will cause the formation of the complex ion cobalt (II) chloride (CoCl_4^{2-}). The resulting solution will have a blue color. The equilibrium between cobalt (II) ion and cobalt (II) chloride ion is as follows:



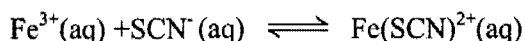
Pink

Blue

Increasing the concentration of chloride ions (from HCl) will increase the concentration of a reagent, causing the position of equilibrium to shift. Whereas, increasing the amount of water will increase the concentration of a product, causing the equilibrium to shift to relieve the stress. A color change will indicate the shift in equilibrium.

Part 3: Equilibrium of the $\text{Fe}(\text{SCN})_2^{2+}$ Ion

A solution of aqueous iron (III) ions, Fe^{3+} , has a yellow color. When thiocyanate ions, SCN^- , are added to the solution, iron (III) thiocyanate ions, $\text{Fe}(\text{SCN})_2^{2+}$, will be formed; and the solution will become red in color. The equilibrium for this reaction is:



Yellow

Red

The addition of Fe^{3+} ions or SCN^- ions will increase the concentration of a reagent, causing the position of equilibrium to shift to relieve the stress. Removing Fe^{3+} ions or SCN^- ions will decrease the concentration of a reagent, also causing the position of equilibrium to shift. The shifts will be obvious as the color of the reaction mixture changes, either becoming more yellow or deeper red. Note that adding OH^- ions to the solution will form the insoluble salt, iron (III) hydroxide, $\text{Fe}(\text{OH})_3(\text{s})$, causing a decrease in the concentration of Fe^{3+} ions in solution.

Buffers

The body's pH must remain fairly constant; a significant increase or decrease in pH can cause considerable damage to the body and possibly death. Buffer systems in the blood help to keep the body at a constant pH. A **buffer** is an aqueous

solution that is able to resist pH changes when acids or bases are added to it. Common buffers contain equal amounts of a weak acid and its conjugate base. The acid can react with OH^- ions from bases; the conjugate base can react with H_3O^+ ions from acids. Adding small quantities of acids or bases to a buffer will not have a large effect on the original pH of the buffer.

The amount of acid or base a buffer can accommodate depends upon the **buffer capacity**. A buffer can absorb only a certain amount of H_3O^+ ions or OH^- ions and still maintain a constant pH. Eventually, the buffer capacity will be exceeded, and the buffer will stop working. Buffers will have different buffer capacities depending on the concentrations of acid and conjugate base initially present.

The **Henderson-Hasselbalch equation** is used to find the pH of buffer solutions. The equation is:

$$\text{pH} = \text{pK}_a + \log \frac{[\text{A}^-]}{[\text{HA}]}$$

[HA] refers to the concentration of the weak acid, such as $\text{HC}_2\text{H}_3\text{O}_2$. $[\text{A}^-]$ refers to the concentration of the conjugate base, such as $\text{C}_2\text{H}_3\text{O}_2^-$ from $\text{NaC}_2\text{H}_3\text{O}_2$. The pK_a is equal to the $-\log$ of the K_a . K_a and pK_a values can be found in Table 9.9 on page 298 of your textbook.

This lab explores how buffer solutions differ from water when acids or bases are added by monitoring changes in pH.

Parts 1 & 2: Addition of Acids and Bases

A buffer system will be tested against water. Changes in the pH of the solution, due to the addition of HCl or NaOH, will be monitored using the pH probe.

Part 3: Buffer Capacity

A new buffer solution will be prepared, having more conjugate base $\text{C}_2\text{H}_3\text{O}_2^-$ than the original buffer. Changes in its pH due to the addition of HCl will be monitored to investigate the concept of buffer capacity.

Procedure

Le Chatelier's Principle

Part 1: Equilibrium of a Saturated NaCl Solution

- A. Label two 4" test tubes **1B** and **1C**.
- B. Add 5 mL of the saturated sodium chloride (NaCl) solution to test tubes **1B** and **1C**. Describe the appearance of this solution on line **B** of Data Sheet I.

Caution: HCl is corrosive and toxic. Prevent contact with your skin, clothing, and eyes. Do not ingest.

- C. Add several drops of 12 M hydrochloric acid (HCl) to the solution in test tube **1C**. Describe its appearance on line **C** of Data Sheet I.

Part 2: Equilibrium of $\text{Co}(\text{H}_2\text{O})_6^{2+}$ and CoCl_4^{2-}

- D. Label two 4" test tubes **2E** and **2F**.
- E. Add 5 drops of the 1 M cobalt (II) chloride solution [$\text{CoCl}_2(\text{aq})$] to test tube **2E** and **2F**. Describe the appearance of the solution on line **E** of Data Sheet I.

Caution: HCl is corrosive and toxic. Prevent contact with your skin, clothing, and eyes. Do not ingest.

- F. Add 12 M HCl dropwise to test tube **2F**, tapping to mix after each drop, until the solution has an obvious change in color. Describe its appearance on line **F** of Data Sheet I.
- G. Add distilled water dropwise to test tube **2F**, tapping to mix after each drop, until the solution has an obvious change in color. Describe its appearance on line **G** of Data Sheet I.

Part 3: Equilibrium of the $\text{Fe}(\text{SCN})^{2+}$ Ion

- H. Connect the colorimeter to the CH1 port of the LabPro interface. Calibrate the colorimeter according to the calibration rules in the front of the manual. During the second part of the calibration, turn the wavelength knob to *blue*.

- I. Label four 4" test tubes, **3J**, **3K**, **3L**, and **3M**.

- J. Add 5 mL of the iron (III) thiocyanate solution [$\text{Fe}(\text{SCN})^{2+}(\text{aq})$] to each of the four test tubes, **3J**, **3K**, **3L**, and **3M**. Describe the appearance of the solution on line **J** of Data Sheet I. Using the technique for data collection described in the calibration rules, find the absorbance of the solution, and record it on line **J** of Data Sheet I.

- K. Add 20 drops of 0.1 M iron (III) chloride solution [$\text{FeCl}_3(\text{aq})$] to test tube **3K**. Tap the test tube to mix. Describe its appearance and record its absorbance on line **K** of Data Sheet I.

- L. Add 20 drops of 0.1 M potassium thiocyanate solution [$\text{KSCN}(\text{aq})$] to test tube **3L**. Tap the test tube to mix. Describe its appearance and record its absorbance on line **L** of Data Sheet I.

Caution: NaOH is corrosive and toxic. Prevent contact with your skin, clothing, and eyes. Do not ingest.

- M. Add 5 drops of 6 M NaOH, to test tube **3M**. Tap the test tube to mix. Describe its appearance and record its absorbance on line **M** of Data Sheet I.

Buffers

- A. Label four 100 mL beakers, **1**, **2**, **3**, and **4**.
- B. Add 50 mL of distilled water to beakers **1** and **3**.
- C. Add 50 mL of 1.0 M acetic acid ($\text{HC}_2\text{H}_3\text{O}_2$) to beakers **2** and **4**. Then completely dissolve 4.1 g of sodium acetate ($\text{NaC}_2\text{H}_3\text{O}_2$) in each solution. Label both beakers as **Buffer 1**.

Part 1: Addition of Acid

- D. Connect the pH probe into the CH2 port of the LabPro interface. Calibrate the pH probe according to the calibration rules for the pH probe.
- E. Using the pH probe, find the pH of the distilled water in beaker 1. Record the pH on Data Sheet II. Leave the probe in the water.

Caution: HCl is corrosive and toxic. Prevent contact with your skin, clothing, and eyes. Do not ingest.

- F. Obtain about 30 mL of 6 M HCl in a labeled 50 mL beaker.
- G. Using a disposable pipette, add 1.5 mL of 6 M HCl to the water in beaker 1. Stir with the probe and record the new pH on Data Sheet II. Keep adding HCl in 1.5 mL increments until a total of 9 mL has been added. After each addition, record the pH on Data Sheet II.
- H. Measure the pH of Buffer 1 in beaker 2 using the pH probe. Record the pH on Data Sheet II. Leave the probe in the buffer.
- I. Using a disposable pipette, add 1.5 mL of 6 M HCl to Buffer 1 in beaker 2. Stir with the probe and record the new pH on Data Sheet II. Keep adding HCl in 1.5 mL increments until a total of 9 mL has been added. After each addition, record the pH on Data Sheet II.

Part 2: Addition of Base

- J. Using the pH probe, find the pH of the distilled water in beaker 3. Record the pH on Data Sheet II. Leave the probe in the water.

Caution: NaOH is corrosive and toxic. Prevent contact with your skin, clothing, and eyes. Do not ingest.

- K. Obtain about 30 mL of 6 M NaOH in a labeled 50 mL beaker.
- L. Using a disposable pipette, add 1.5 mL of 6 M NaOH to the water in beaker 3. Stir with the probe and record the new pH on Data Sheet II. Keep adding NaOH in 1.5 mL increments until a total of 9 mL has been added. After each addition, record the pH on Data Sheet II.
- M. Find the pH of Buffer 1 in beaker 4 using the pH probe. Record the pH on Data Sheet II. Leave the probe in the buffer.
- N. Using a disposable pipette, add 1.5 mL of 6 M NaOH to Buffer 1 in beaker 4. Stir with the probe and record the new pH on Data Sheet II. Keep adding NaOH in 1.5 mL increments until a total of 9 mL has been added. After each addition, record the pH on Data Sheet II.

Part 3: Buffer Capacity

- O. Obtain a 100 mL beaker. Label it **Buffer 2**. Add 50 mL of 1.0 M acetic acid ($\text{HC}_2\text{H}_3\text{O}_2$). Then, completely dissolve 10 g of sodium acetate ($\text{NaC}_2\text{H}_3\text{O}_2$) in the solution.
- P. Using the pH probe, measure the pH of **Buffer 2**. Record the pH on Data Sheet II. Leave the probe in the buffer.

Caution: HCl is corrosive and toxic. Prevent contact with your skin, clothing, and eyes. Do not ingest.

- Q. Using a disposable pipette, add 1.5 mL of 6 M HCl to **Buffer 2**. Stir with the probe and record the new pH on Data Sheet II. Keep adding HCl in 1.5 mL increments until a total of 9 mL has been added. After each addition, record the pH on Data Sheet II.

Post-Lab Questions

1. Finish Table 1. Complete the "Observation" column using your observations recorded on Data Sheet I. For each observation, decide if it follows the prediction made. If so, record that your prediction was correct in the "Interpretation" column. If an observation does not follow the prediction, explain why your prediction was incorrect in the "Interpretation" column.

Buffer capacity is dependent on two factors: initial pH and component concentrations. The buffer capacity is greatest when its initial pH is equal to the pK_a of the weak acid present in the buffer. Also, the greater the concentrations of the weak acid and its conjugate base, the higher the buffer capacity.

3. The pH of Buffer 2 did not drop as drastically as Buffer 1 after the same volume of acid was added to each. Why?
2. Identify the connection between buffer action and Le Chatelier's principle.
4. Explain why it is fortunate to the human body that blood has a good buffering system. What would happen if your blood did not contain a buffering system, but was similar to water?

Name _____

Date _____

Data Sheet I

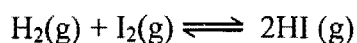
	Reaction	Observation	Absorbance
B	NaCl solution		
C	NaCl solution + HCl		
E	CoCl ₂ solution		
F	CoCl ₂ solution + HCl		
G	CoCl ₂ solution + HCl + H ₂ O		
J	Fe(SCN) ²⁺ solution		
K	Fe(SCN) ²⁺ solution + FeCl ₃		
L	Fe(SCN) ²⁺ solution + KSCN		
M	Fe(SCN) ²⁺ solution + NaOH		

Data Sheet II

Part 1	Initial pH	+ 1.5 mL HCl	+ 3.0 mL HCl	+ 4.5 mL HCl	+ 6.0 mL HCl	+ 7.5 mL HCl	+ 9.0 mL HCl
Beaker 1: Water							
Beaker 2: Buffer 1							
Part 2	Initial pH	+ 1.5 mL NaOH	+ 3.0 mL NaOH	+ 4.5 mL NaOH	+ 6.0 mL NaOH	+ 7.5 mL NaOH	+ 9.0 mL NaOH
Beaker 3: Water							
Beaker 4: Buffer 1							
Part 3	Initial pH	+ 1.5 mL HCl	+ 3.0 mL HCl	+ 4.5 mL HCl	+ 6.0 mL HCl	+ 7.5 mL HCl	+ 9.0 mL HCl
Buffer 2							

Pre-Lab Questions

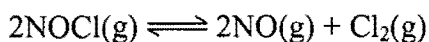
Equilibrium concentrations of reactants and products can be expressed as a constant, K_{eq} . This constant is equal to the concentration of each product, raised to the power of its respective stoichiometric coefficient, divided by the concentration of each reactant, raised to the power of its respective stoichiometric coefficient. Concentrations at equilibrium can be predicted using the equilibrium constant. The equilibrium constant for the reaction



can be written as follows:

$$K_{eq} = \frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]}$$

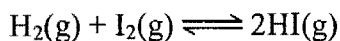
1. The equilibrium constant for the reaction



is 9.9×10^{-2} M at 25°C . The equilibrium concentrations of NOCl and Cl_2 are 5.2 M and 0.51 M respectively. Calculate the equilibrium concentration of NO for this reaction at 25°C .

Temperature changes affect the position of equilibrium. Increasing the temperature of a reaction is equivalent to adding heat. For an endothermic reaction, heat is a reactant; therefore, increasing the temperature would shift the position of equilibrium to the right.

2. The following reaction is exothermic.



Is heat a product or a reactant?

How would the position of equilibrium shift if the reaction was run at a lower temperature?

3. Table 1 includes the reactions from the Le Chatelier's principle portion of this lab. Column one displays all the reactions and added stresses contained in today's lab. In column two, indicate if the position of equilibrium will shift to the left or right due to the added stress. In the example, the position of equilibrium is predicted to shift to the left because the concentration of one of the products, Cl^- , will increase.

In the "Prediction" column, predict what you should observe due to the added stress. As the example shows, an increase in the concentration of Cl^- ions should shift the equilibrium to the left causing solid NaCl to precipitate. The clear, aqueous solution should become cloudy with obvious solid NaCl present.

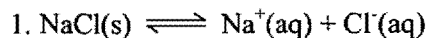
During the lab, record observations on Data Sheet I. Those observations will be used to fill in the remainder of Table 1 during post-lab.

4. In this lab, you will prepare two buffer solutions. Buffer 1 will have 4.1 g of sodium acetate ($\text{NaC}_2\text{H}_3\text{O}_2$) dissolved in 50 mL of 1.0 M acetic acid ($\text{HC}_2\text{H}_3\text{O}_2$). The pK_a of acetic acid is 4.75.

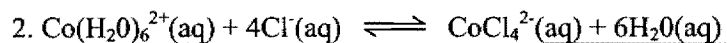
Predict the pH of this buffer system using the Henderson-Hasselbalch equation.

Buffer 2 will have 10 g of sodium acetate dissolved in 50 mL of 1.0 M acetic acid. Predict its pH.

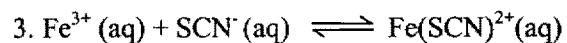
Table 1: Pulling It All Together



Reaction	Shift Left or Right	Prediction	Observation	Interpretation
NaCl solution + HCl	Left	As Cl^- ion concentration increases, NaCl will precipitate out of solution and solid will appear	Solid appeared as HCl was added dropwise	My prediction was correct.



Reaction	Shift Left or Right	Prediction	Observation	Interpretation
CoCl_2 solution + HCl				
CoCl_2 solution + HCl + H_2O				



Reaction	Shift Left or Right	Prediction	Observation	Interpretation
Fe(SCN)^{2+} solution + FeCl_3				
Fe(SCN)^{2+} solution + KSCN				
Fe(SCN)^{2+} solution + NaOH				

Appendix C

General, Organic, and Biochemistry for the Health Sciences

**General, Organic,
AND
Biochemistry
FOR THE
Health Sciences
CHEM 101 LABORATORY MANUAL**

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Ball State University



**Boston Burr Ridge, IL Dubuque, IA Madison, WI New York
San Francisco St. Louis Bangkok Bogotá Caracas Kuala Lumpur
Lisbon London Madrid Mexico City Milan Montreal New Delhi
Santiago Seoul Singapore Sydney Taipei Toronto**

The following students helped us design, plan, test, and write many of the laboratory experiments in this manual and even a few not included in the manual. We could not have completed this project without their contributions.

Amanda K. Brooks; Saponification – The Process of Making Soap

Dalahnna E. Cooper; Preparation and Identification of Esters and others not included in the manual

Jennifer G. Dotson; Density and Chemistry of Copper and Its Quantitative Recovery

Andi Grove; Density; Chemistry of Copper and Its Quantitative Recovery, and others not included in the manual

Sheri Isaacs; Density and others not included in the manual

Andrew Mummert; Enzymes and others not included in the manual

Katie Smitherman; Properties of Solutions and others not included in the manual

Joseph B. Stout; Alkenes and Alcohols and Their Classification

General, Organic, and Biochemistry for the Health Sciences
CHEM 101 LABORATORY MANUAL

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Laboratory Safety

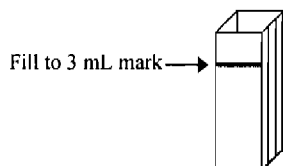
1. Safety goggles must be worn at *all* times.
2. Wearing closed toe shoes and long pants is highly advised.
3. Always work under the supervision of an instructor or with a partner in the lab.
4. Know the location of the following laboratory safety devices:
 - a. Shower
 - b. Eyewash
 - c. Fire blanket
 - d. Fire extinguisher
5. Long hair should be tied back.
 - a. Avoid direct contact with any chemical, especially on your hands and / or face.
6. Use the “whaff” procedure to smell odors.
7. Chemicals should not be ingested.
8. No food, drink, or gum is allowed in the lab.
9. Keep each lab station clean; avoid clutter.
10. Discard wastes appropriately according to the instructor.
11. Keep all walkways, emergency equipment, and exit routes clear.
12. Inspect glassware before using it. Do not use glassware with cracks and scratches.
13. Report broken glassware, chemical spills, or unsafe conditions to the instructor.
14. Be certain that gas and hotplates are turned off before you leave.

Always inform the instructor of damage to equipment, spills, or unsafe conditions. This action could prevent an emergency and save a life.

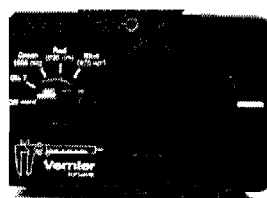
Colorimeter

Calibration

1. Connect the colorimeter to the CH1 port of the LabPro interface.
2. Open the LoggerPro program by clicking the "Current 101" icon on the computer. Click on "Experiment" at the top of the screen. Scroll down to "Calibrate," and select it.
3. Under the "Calibrate" tab, make sure CH1 is selected. Click the "Perform Now" button on the right side of the box. A window called "Reading 1" will appear on the screen near the bottom.
4. By only touching the ribbed sides, obtain a cuvette. Never touch the smooth sides of a cuvette. Fill the cuvette with the blanking solution determined by your instructor to the fill line on the side of the cuvette (3 mL volume, Figure 1). Wipe any excess liquid off the outside of the cuvette with a Kimwipe.
5. Place the cuvette into the cuvette slot of the colorimeter; the fill line of the cuvette should face the white line on the colorimeter (Figure 2).
6. Close the lid. Turn the wavelength knob on the colorimeter to "0% T."
7. The reading in the "Input 1" box of "Reading 1" on the computer will change. Once it has stabilized, enter a "0" into the "Value 1" box. Click "Keep." A new window called, "Reading 2" will appear.
8. Turn the wavelength knob on the colorimeter to desired wavelength.
9. The reading in the "Input 1" box of "Reading 2" on the computer will change. Once it has stabilized, enter "100" into the "Value 2" box. Click "Keep."
10. Click "OK." The colorimeter is now calibrated.



5. Place the cuvette into the cuvette slot of the colorimeter; the fill line of the cuvette should face the white line on the colorimeter (Figure 2).



← Fill line
faces

Data Collection

1. After calibration, empty the cuvette, and rinse with distilled water.
2. To find the absorbance of a solution, fill the cuvette with the solution to the mark on the side of the cuvette. Wipe off excess liquid with a Kimwipe.
3. Place cuvette in cuvette slot, as in step 5 of the calibration.
4. The wavelength knob should remain on the same wavelength used for calibration.
5. An absorbance reading will appear on the screen. Wait until it stabilizes to record.
6. When taking the absorbance of more than one sample, rinse the cuvette with distilled water before filling with a different solution.

pH Probe

Calibration

1. Obtain about 20 mL each of pH 4 buffer and pH 7 buffer in two labeled beakers.
2. Connect the pH probe to the CH2 port of the LabPro interface.
3. Open the LoggerPro program by clicking the "Current 101" icon on the computer.
4. Click on the green "LabPro" icon located on the toolbar at the top of the screen. A box named "Sensor Properties" will appear.
5. Within the "Sensor Setup" tab, make sure CH2 is selected among the icons across the top. To do this, just click the icon above "CH2."
6. In the "Sensor" box, make sure "pH Probe" is selected. Check that the "Calibration" box says "PH."
7. Switch to the "Calibrate" tab by clicking it at the top of the inset box.
8. Click the button, "Perform Now." A window called "Reading 1" will appear near the bottom.
9. Carefully unscrew the cap from the pH probe. Place the solution-filled cap aside without spilling.
10. Rinse pH probe with distilled water. Wipe dry with a Kimwipe. Place in the pH 4 buffer. Watch the value in the "Input 2" window until it is fairly stable. This may take 20-30 seconds.
11. When the value has stabilized, enter a "4" into the "Value 1" box. Click "Keep." A new window called "Reading 2" will appear now.
12. Remove the pH probe and rinse with distilled water. Wipe dry with a Kimwipe. Place in the pH 7 buffer. Watch the value in the "Input 2" window until it is fairly stable.
13. When the value has stabilized, enter a "7" into the "Value 2" box. Click "Keep."
14. Click "OK." This means the pH probe is now calibrated. Keep the pH probe in pH 7 buffer until it is needed again.

Data Collection

1. When it is time to use the pH probe, remove it from the pH 7 buffer.
2. Whenever the probe is transferred from one solution to another, even if it is water, it must always be rinsed well with distilled water.
3. After rinsing, wipe with a Kimwipe.
4. When the probe is in a solution, a pH reading will appear on the screen. Wait until it stabilizes to record.
5. To temporarily store the pH probe, it can be placed in the pH 7 buffer or a beaker of distilled water.
6. When finished with the pH probe, rinse well, and screw the cap back on carefully.

Common Ions

Cations

+1 cations

Ammonium	NH_4^+
Copper (I)	Cu^+
Lithium	Li^+
Potassium	K^+
Silver	Ag^+
Sodium	Na^+

+2 cations

Barium	Ba^{2+}
Beryllium	Be^{2+}
Calcium	Ca^{2+}
Cobalt (II)	Co^{2+}
Copper (II)	Cu^{2+}
Lead (II)	Pb^{2+}
Magnesium	Mg^{2+}
Zinc	Zn^{2+}

+3 cations

Aluminum	Al^{3+}
Chromium (III)	Cr^{3+}
Iron (III)	Fe^{3+}

Anions

-1 anions

Bicarbonate	HCO_3^-
Bromide	Br^-
Chlorate	ClO_3^-
Chloride	Cl^-
Chlorite	ClO_2^-
Cyanide	CN^-
Fluoride	F^-
Hydroxide	OH^-
Iodate	IO_3^-
Iodide	I^-
Nitrate	NO_3^-
Nitrite	NO_2^-
Permanganate	MnO_4^-

-2 anions

Carbonate	CO_3^{2-}
Chromate	CrO_4^{2-}
Dichromate	$\text{Cr}_2\text{O}_7^{2-}$
Oxalate	$\text{C}_2\text{O}_4^{2-}$
Sulfate	SO_4^{2-}
Sulfite	SO_3^{2-}

-3 anions

Phosphate	PO_4^{3-}
-----------	--------------------

Periodic Table of the Elements

1 IA 1 H 1.008	2 IIA 4 Be 9.01											3 IIIA 5 B 10.81	4 IVA 6 C 12.01	5 VA 7 N 14.01	6 VIA 8 O 16.00	7 VIIA 9 F 19.00	8 VIIIA 2 He 4.00
11 Na 22.99	12 Mg 24.31											13 Al 26.98	14 Si 28.09	15 P 30.97	16 S 32.07	17 Cl 35.45	18 Ar 39.95
19 K 39.10	20 Ca 40.08	21 Sc 44.96	22 Ti 47.88	23 V 50.94	24 Cr 52.00	25 Mn 54.94	26 Fe 55.85	27 Co 58.93	28 Ni 58.69	29 Cu 63.55	30 Zn 65.39	31 Ga 69.72	32 Ge 72.61	33 As 74.92	34 Se 78.96	35 Br 79.90	36 Kr 83.80
37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.94	43 Tc (98)	44 Ru 101.07	45 Rh 102.91	46 Pd 106.42	47 Ag 107.87	48 Cd 112.41	49 In 114.82	50 Sn 118.71	51 Sb 121.75	52 Te 127.60	53 I 126.90	54 Xe 131.29
55 Cs 132.91	56 Ba 137.33	57 La 138.91	72 Hf 178.49	73 Ta 180.95	74 W 183.85	75 Re 186.21	76 Os 190.2	77 Ir 192.22	78 Pt 195.08	79 Au 196.97	80 Hg 200.59	81 Tl 204.38	82 Pb 207.2	83 Bi 208.98	84 Po (209)	85 At (210)	86 Rn (222)
87 Fr (223)	88 Ra 226.03	89 Ac 227.03	104 Rf (261)	105 Db (262)	106 Sg (266)	107 Bh (264)	108 Hs (269)	109 Mt (268)	110 Ds (271)	111 Rg (272)	112 Uub (285)	113 Uut (284)	114 Uuq (289)	115 Uup (288)	116 Uuh (292)		

58 Ce 140.12	59 Pr 140.91	60 Nd 144.24	61 Pm (145)	62 Sm 150.36	63 Eu 151.97	64 Gd 157.25	65 Tb 158.93	66 Dy 162.5	67 Ho 164.93	68 Er 167.26	69 Tm 168.93	70 Yb 173.04	71 Lu 174.97
90 Th 232.04	91 Pa 231.04	92 U 238.03	93 Np 237.05	94 Pu (244)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (251)	99 Es (252)	100 Fm (257)	101 Md (258)	102 No (259)	103 Lr (262)

Density

Purpose of the experiment:

Compare a pipette, beaker, and graduated cylinder to determine which volumetric glassware is most accurate. Use the density equation to identify an unknown metal.

Background

Chapter one of your textbook defines **mass** as the quantity of matter in an object, and **volume** as space occupied by a liquid, solid or gas. Together, these two measurements define density. The **density** (d) of any substance is defined as its mass (m) per unit volume (V).

$$d = m / V$$

Liquids, solids, and gases all have densities.

Density is an example of a physical property. We can use physical properties to identify an unknown substance. The density of a substance is a physical property that is dependent upon temperature. This means that the density of a substance will change when the temperature changes. This occurs because the volume of solids, liquids, and gases will change when the temperature changes; whereas, the mass of a substance will remain unchanged as the temperature changes. For most substances, the volume will increase as the temperature increases. Thus, as the temperature increases, the density of most substances will decrease. One notable exception to this is water. Starting at 4°C, the density of water decreases as the temperature is decreased. At 0°C, ice is less dense than water, explaining why ice floats in water.

Density is calculated by dividing the mass of a substance by its volume. Thus, to calculate the density, both the mass and volume of a substance must be determined. A balance is used to determine the mass of a substance. Volumes of solids, liquids, and gases are measured in different ways. The volume of a gas is simply equal to the volume of its container. Volumetric glassware is typically used to measure the volume of liquids. Volumes of solids cannot be measured so easily. For a regularly shaped solid, the volume can be calculated from the dimensions of the object, for example, the volume of a rectangular block can be determined by multiplying the length, width, and height together. For an irregularly shaped solid, the volume can be determined by placing the substance in a known volume of water and

measuring the **displacement** of the water. Displacement is an indicator of the volume of an object, for when placed in a liquid, the object will move a volume of liquid equal to its own volume. The volume of a quarter can be determined using the displacement method. First a known volume of water, such as 25.0 mL, would be added to a graduated cylinder. Then, the quarter would be added to the water, and a certain volume of water would be moved, or displaced. The new volume of water would need to be recorded. If the new volume of water was 26.5 mL, the volume of water displaced was 1.5 mL. And, the volume of the quarter would be 1.5 mL. To determine the density, simply divide the mass of the quarter by the volume just determined.

Today, you will be determining the density of an unknown metal using the same two steps described above. After finding the mass of the unknown metal, you will determine the volume of the metal using the water displacement method. The density of the metal then can be determined. By comparing your calculated density to the density of known metals, you may identify the unknown metal.

Because you will be using the balance in the lab, it is important to use it correctly. Always follow these directions when using the balance:

1. Press the **On** button to turn on the balance.
2. Place the container in which the substance will be weighed on the balance.
3. Zero the mass of the empty container by pressing the **O/T** (tare) button.
4. Wait for the screen to read 0.00.
5. Add the substance to the container until the desired mass is obtained.
6. Wait for the balance to stabilize before recording the mass.
7. Keep the balance clean by brushing off any spills with the balance brush.

Procedure

I. Volumetric glassware accuracy

- A. Obtain approximately 150 mL of distilled water in a 250-mL beaker. Record the temperature on Data Sheet I. See the instructor for the density of water at this temperature, and record it on Data Sheet I.
- B. Weigh a dry 50-mL beaker using a balance. Record the mass on line **B** of Data Sheet I to two decimal places.
- C. Obtain a 25-mL pipette. Fill the pipette up to the marked line with distilled water from the 250-mL beaker. Notice that the water curves; it is lower in the center than on the sides. This feature is called a meniscus. The lowest point of the meniscus should be used to find the volume. Record the volume on line **C** of Data Sheet I.
- D. Transfer the distilled water from the pipette into the pre-weighed 50-mL beaker.
- E. Find the mass of the beaker with the water. Record the mass of the beaker and water on line **E** of Data Sheet I to two decimal places.
- F. Weigh a **dry** 50-mL beaker using a balance. Record the mass on line **F** of Data Sheet I to two decimal places.
- G. Transfer 30 mL of distilled water from the 250-mL beaker to the 50-mL beaker. Use the markings on the 50-mL beaker to know when to stop. Record the volume on line **G** of Data Sheet I.
- H. Find the mass of the beaker containing the 30 mL of distilled water. Record this mass on line **H** of Data Sheet I to two decimal places.
- I. Weigh a 50-mL graduated cylinder using a balance. Record the mass on line **I** of Data Sheet I to two decimal places.
- J. Transfer 30 mL of distilled water from the 250-mL beaker to the graduated cylinder. To obtain an accurate volume, be certain to record the volume using the bottom of the meniscus. Record the volume on line **J** of Data Sheet I.

- K. Weigh the graduated cylinder containing the water. Record the mass on line **K** of Data Sheet I to two decimal places.

II. Density of an unknown metal

- L. Obtain an unknown metal sample from the laboratory instructor. Record the identifying number of your sample on line **L** of Data Sheet II.

<p>Note: Do not touch the metal sample with your fingers. Fingerprints will cause an erroneous mass reading.</p>
--

- M. Using a dampened paper towel, carefully wipe the metal sample clean. Then, dry the sample thoroughly with a dry paper towel. Obtain approximately 30 g of the metal in a weighing boat. Record the mass of the metal sample on line **M** of Data Sheet II to two decimal places.
- N. Add 30 mL of water to the 50-mL graduated cylinder. Read the volume of liquid in the graduated cylinder to the nearest 0.1 mL. Record this volume on line **N** of Data Sheet II.
- O. Slightly tilt the graduated cylinder and transfer the metal sample into the graduated cylinder. Avoid splashing any water out of the cylinder.
- P. Read the volume of the liquid in the graduated cylinder to the nearest 0.1 mL. Record this volume on line **P** of Data Sheet II.
- Q. Drain the water from the graduated cylinder being careful not to lose any of the metal down the sink. Collect the metal sample on a paper towel and dry. Return the metal sample to the laboratory instructor.

III. Calculations

Volumetric Glassware Accuracy

- R.** Calculate the density of the water in the pipette by dividing the mass of water (**E-B**) by the volume of water (**C**). Record the density on line **R** of Data Sheet I to the correct number of significant figures.
- S.** Calculate the density of the water in the beaker by dividing the mass of water (**H-F**) by the volume of water (**G**). Record the density on line **S** of Data Sheet I to the correct number of significant figures.
- T.** Calculate the density of the water in the graduated cylinder by dividing the mass of water (**K-I**) by the volume of water (**J**). Record the density on line **T** of Data Sheet I to the correct number of significant figures.

- U.** Find the percent error of each density measurement using the equation on Data Sheet I.

Density of an unknown metal

- V.** Calculate the volume of the metal sample by finding the volume of the displaced water (**P-N**). Record the volume on line **V** of Data Sheet II.
- W.** Calculate the density of the unknown metal by dividing the mass of the metal by its volume. (**M/V**). Record the density on line **W** of Data Sheet II to the correct number of significant figures.
- X.** Use the given list of metals and their densities to determine the identity of your unknown metal. Record the metal sample identity on line **X** of Data Sheet II.

Name _____

Date_____

Post-Lab Questions

- Which volumetric glassware did you find to be the most accurate at measuring a volume of water?
 - Can you trust the volume markings on a beaker?
- The density of titanium is 4.54 g/mL. What is the mass, in g, of 17.3 mL of titanium?
- Ships have hulls constructed of steel, an iron alloy, which has a density around 7.8 g/mL. Explain how it is possible for ships to float on the surface of water when they are constructed of a substance which is more dense than water.
- An amateur gold miner discovered some yellow nuggets which had a mass of 105 g and a volume of 21 cm³. Are the nuggets gold or pyrite, fool's gold? The density of gold is 19.3 g/mL. Pyrite has a density of 5.0 g/mL. Show your work to justify your answer.

Name _____

Date _____

Data Sheet I**Density Equation:**

$$\text{density} = \frac{\text{mass (g)}}{\text{Volume (mL)}}$$

Percent Error Equation:

$$\frac{(\text{accepted density} - \text{calculated density})}{(\text{accepted density})} \bullet 100\%$$

A. Temperature of water: _____ °C

Accepted density of water: _____ g/mL

	25 mL Pipette	50 mL Beaker	50 mL Graduated Cylinder
Mass of glassware	B.	F.	I.
Volume of water	C.	G.	J.
Mass of glassware + water	E.	H.	K.
Mass of water	(E-B)	(H-F)	(K-I)
Calculated density of water	R.	S.	T.
% Error			

Name _____

Date _____

Data Sheet II

L. Unknown metal sample number _____

M. Mass of unknown metal _____ g

N. Original volume of water _____ mL

P. Volume of water after adding metal _____ mL

V. Volume of unknown metal (P-N) _____ mL

W. Density of unknown metal _____ g/mL

X. Identification of unknown metal _____

Pre-Lab Questions

Density is calculated by dividing the mass of a substance by its volume, using the equation

$$d = m / V$$

where d is the density, m is the mass, and V is the volume.

Example 1

If 73.2 mL of a liquid has a mass of 61.5 g, what is its density in g/mL?

Solution: We are given the mass and volume values. When inserting the values into the equation above, we get:

$$d = 61.5 \text{ g} / 73.2 \text{ mL}$$

The density of this liquid is 0.840 g/mL.

Example 2

The density of iron is 7.86 g/mL. What is the volume, in mL, given that the piece of iron has a mass of 524 g?

Solution: We are given the mass and the density. We can rearrange the density equation to solve for the volume. Multiplying both sides of the density equation by the volume gives

$$d V = m$$

Dividing both sides by the density, we obtain

$$V = m / d$$

Now, we can solve for the volume by dividing the mass by the density.

$$V = 524 \text{ g} / 7.86 \text{ g/mL}$$

The volume of the piece of iron is 66.7 mL.

Example 3

Percent error is an indicator of the accuracy of a person's results compared to accepted values. To determine percent error, use the equation:

$$\frac{(\text{accepted density} - \text{calculated density})}{(\text{accepted density})} \bullet 100\%$$

After doing an experiment, a student calculated the density of a sample of methanol at 20 °C to be 0.670 g/mL. The student then found the accepted density for methanol at 20 °C in a reference book to be 0.791 g/mL. What is the student's percent error for the density of methanol?

Solution: We are given the calculated density and accepted density values. We know the percent error equation is:

$$\frac{(\text{accepted density} - \text{calculated density})}{(\text{accepted density})} \bullet 100\%$$

When inserting the values, it looks like:

$$\frac{0.791 \text{ g/mL} - 0.670 \text{ g/mL}}{0.791 \text{ g/mL}} \bullet 100\%$$

The student's percent error is: 15.3%, meaning she was only 84.7% accurate.

Name _____

Date _____

Problems

1. A student found the mass of a 50-mL beaker to be 29.656 g. Express this quantity in kilograms and milligrams.
2. An unknown substance has a mass of 56.8 g and occupies a volume of 23.4 mL. What is its density in g/mL?
3. A brain weighing 3.0000 pounds (1360.8 g) occupies a volume of 0.6200 L. What is the density of the brain in g/mL?
4. You are given a piece of metal of unknown origin by a friend. Although the identity of this metal is unknown, your friend believes that the metal may be aluminum. The mass of the metal was found to be 8.91 g. What volume of water would the metal displace if it is aluminum? Aluminum metal has a density of 2.70 g/mL.
5. Using the water displacement method, a student found the density of copper metal to be 8.51 g/mL at 20 °C. The accepted density for copper at 20 °C is 8.92 g/mL. Calculate the student's percent error.

The Chemical Formula of an Oxide of Magnesium

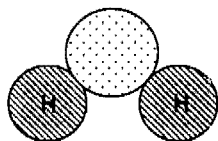
Purpose of the Experiment:

Determine the chemical formula of an oxide of magnesium.

Background

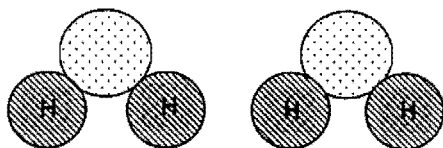
When two or more different atoms combine to form a pure substance, it is called a compound. There are two types of compounds, molecules and ionic compounds. You will learn more about each of these as the semester progresses. The composition of either type of compound is symbolized by its chemical formula. A chemical formula includes the atomic symbol of each element present in the compound followed by a subscript specifying the number of atoms present. If a subscript is absent, only one atom of that element is present in the compound.

A water molecule contains two atoms of hydrogen and one atom of oxygen. Thus, water has the chemical formula, H_2O .



One water molecule contains two hydrogen atoms and one oxygen atom.

The subscripts act as an atom ratio. In each water molecule, there is a 2:1 ratio of hydrogen atoms to oxygen atoms. If we were to increase the number of water molecules, the ratio would still remain at 2:1, even though the total number of atoms changes. For example, two water molecules would consist of four hydrogen atoms and two oxygen atoms, ten water molecules would consist of twenty hydrogen atoms and ten oxygen atoms, 1000 water molecules would consist of 2000 hydrogen atoms and 1000 oxygen atoms, and 6.022×10^{23} water molecules would consist of 12.044×10^{23} hydrogen atoms and 6.022×10^{23} oxygen atoms.



Two water molecules contain four hydrogen atoms and two oxygen atoms.

The value 6.022×10^{23} is called Avogadro's number and is often referred to as a **mole**. Just as there are twelve doughnuts in one dozen of doughnuts, there are 6.022×10^{23} atoms in one mole of atoms. Thus, we can rewrite " 6.022×10^{23} water molecules would consist of 12.044×10^{23} hydrogen atoms and 6.022×10^{23} oxygen atoms" as "*one mole* of water molecules would consist of *two moles* of hydrogen atoms and *one mole* of oxygen atoms". Because a water molecule has an atom ratio of 2:1 (two hydrogen atoms for every one oxygen atom), it also has a mole ratio of 2:1 (two moles of hydrogen for every one mole of oxygen). The subscripts in a chemical formula not only give the atom ratio of the atoms in the compound, they also give the **mole ratio** of atoms in the compound. Chemical formulas represent the mole ratios of the atoms in a compound.

The mole is very useful because it also relates the number of atoms in a sample (a number too large for you to count in your lifetime) to the mass of those atoms in grams (it only takes a few seconds to weigh a sample on a balance). The mole can be defined as the number of atoms present in a 12.01 g sample of carbon. A 12.01 g sample of carbon contains 6.022×10^{23} carbon atoms, or one mole of atoms. Thus, one mole of carbon weighs 12.01 g. The mass of one mole of a substance is called a **molar mass**. The molar mass of most elements is equal to its atomic mass expressed in grams. The atomic mass of calcium is 40.08 amu; thus, one calcium atom weighs 40.08 amu. By definition, the molar mass of calcium is 40.08 g/mol. This means that a 40.08 g sample of calcium contains one mole, or 6.022×10^{23} atoms.

The molar mass can be used to interconvert between the number of moles and the mass of an element. To find the mass of a certain number of moles of an element, simply multiply the number of moles by the element's molar mass. Because 1.0 mole of calcium has a mass of 40.08 g, 1.5 moles of calcium would have a mass of 60.12 g ($40.08 \text{ g/mol} \times 1.5 \text{ mol}$). To find the number of moles of an element from its mass, simply divide the mass by the molar mass. Because a 40.08 g

sample of calcium consists of 1.0 mole of calcium, 80.16 g of calcium would contain 2.0 moles (80.16 g ÷ 40.08 g/mol).

The molar mass of a compound can be determined from its chemical formula because the chemical formula of a compound represents the mole ratio of atoms in that compound. One mole of water molecules consists of two moles of hydrogen atoms and one mole of oxygen atoms. The mass of two moles of hydrogen is 2.02 g, and the mass of one mole of oxygen is 16.00 g. Thus, the mass of one mole of water is 18.02 g, making the molar mass of water 18.02 g/mol.

Chemical formulas do not, however, represent the mass ratio of atoms in a compound. The mass ratio of hydrogen to oxygen is 1:8; whereas, the mole ratio is 2:1. Because a chemical formula represents a mole ratio, it is necessary to convert any masses determined in the lab to moles. Then, a mole ratio can be determined. Using the mole ratio, a chemical formula can be assigned.

Take the following as an example. A student heated 0.70 g of nitrogen in an open container and produced a compound of nitrogen and oxygen. The compound of nitrogen and oxygen had a mass of 2.3 g. From this information, the chemical formula can be determined using the following steps.

1. To begin, the mass of each element in the compound must be determined. Based on the information given, there are 0.70 g of nitrogen present. Because only one other element is present in the compound, the mass of oxygen can be determined by subtracting the mass of nitrogen from the mass of the compound.

$$2.3 \text{ grams} - 0.7 \text{ grams} = 1.60 \text{ grams oxygen}$$

2. To determine the number of moles of nitrogen and oxygen in the compound, divide the mass of each element by their respective molar mass.

$$\text{moles of N} = 0.70 \text{ g N} \left(\frac{1 \text{ mol N}}{14.01 \text{ g N}} \right) = 0.050 \text{ mol N}$$

$$\text{moles of O} = 1.60 \text{ g O} \left(\frac{1 \text{ mol O}}{16.00 \text{ g O}} \right) = 0.100 \text{ g O}$$

3. To calculate the mole ratio, the number of moles of each element must be divided by the smaller number of moles. In this example, the 0.050 moles of nitrogen is used because it is the smaller number of moles.

$$\text{N ratio} = \frac{\# \text{ mol N}}{\# \text{ mol N}} = \frac{0.050 \text{ mol N}}{0.050 \text{ mol N}} = 1.0$$

$$\text{O ratio} = \frac{\# \text{ mol O}}{\# \text{ mol N}} = \frac{0.100 \text{ mol O}}{0.050 \text{ mol N}} = 2.0$$

There are two moles of oxygen for every one mole of nitrogen present in this compound. Therefore, the chemical formula would be NO_2 .

Sometimes mole ratios do not contain whole numbers. Many compounds have ratios such as 1:1.5. In these cases, ratios must be multiplied by a common factor in order to get two whole numbered subscripts. For the ratio 1:1.5, if both numbers are multiplied by 2, the ratio becomes 2:3. Table 1 lists some common ratios and their respective whole number ratios.

Ratio	Common Factor	Whole Number Ratios
1:1.67	3	3:5
1:1.50	2	2:3
1:1.33	3	3:4
1:1.25	4	4:5

Table 1

In this lab, you will heat a magnesium sample to form a compound of magnesium and oxygen. Using calculations like those outlined above, you will determine the chemical formula of the oxide of magnesium.

Procedure

- A. Attach an iron ring to a ring stand, allowing enough room underneath for a Bunsen burner. Place a clay triangle on the ring.

Note: Always use a clean, dry crucible. If they are not clean, masses will not be accurate, and results will contain error. Always use crucible tongs to handle the crucible and cover to prevent the transfer of fingerprints. **Never touch a hot crucible.**

- B. Place a porcelain crucible at a slight angle on the triangle. Place the crucible cover on the crucible so that it is slightly ajar. See **Figure 1**.



Figure 1

- C. Turn on the gas and light the Bunsen burner. Adjust burner to produce a very hot flame 1-2 inches in height.
- D. Heat the crucible gently for 5 minutes by moving the flame back and forth across the bottom of the crucible. After 5 minutes, turn off the gas to extinguish the flame.
- E. Remove the cover from the crucible. Set it on wire gauze to cool. Next, remove the crucible, and allow it to cool on the wire gauze.

Note: Never weigh a hot object, for the mass will not be accurate. To determine if an object is cool enough, hold your hand about 2 cm from the object. If no heat is felt on your hand, it is cool enough to weigh.

- F. When the crucible is cool, use crucible tongs to transfer the crucible and cover to the balance. Determine the mass of both together. Record this mass on line **F** of Data Sheet I.

- G. Obtain 3 strips of magnesium ribbon. Loosely roll the magnesium ribbon into a ball, and place it in the crucible.

- H. Weigh the crucible, magnesium, and cover together on the balance. Record this mass on line **H** of Data Sheet I.

- I. Place the crucible with magnesium at a slight angle on the triangle with the cover slightly ajar, as in **Figure 1**.

- J. Turn on the gas and light the Bunsen burner.

Caution: Never look directly into the white flame of burning magnesium, for Mg emits UV light when burned. Eye damage could result.

Note: Completely cover the crucible at the first sign of white smoke. This white smoke is the product, the oxide of magnesium. Allowing it to escape will cause error in your results.

- K. Gently heat the crucible by slowly moving the flame back and forth across the bottom of the crucible. If white smoke is produced, remove the flame; and cover the crucible **completely**.
- L. After about 15-20 seconds, lift the cover slightly, and check to see if white smoke is still present. If smoke is present, repeat steps J through L until you no longer observe smoke. You may have to repeat these steps many times. Proceed to step M when no smoke appears.
- M. Place the cover so that it is slightly ajar, as in **Figure 1**. Light the Bunsen burner, and place it under the crucible so that the flame contacts the lower edge of the crucible. Heat the crucible with a strong flame until the bottom of the crucible is red. Heat for 10 minutes.
- N. Turn off the gas to extinguish the flame.
- O. Remove the cover, invert it, and allow it to cool on the wire gauze. Let the crucible cool on the triangle until no heat is felt when you hold your hand about 2 cm away.

Caution: Before adding the water in step Q, let the crucible cool to room temperature. Adding water to a warm crucible will cause it to shatter and could result in serious injury.

- P. Add 10 drops of deionized water to the crucible.
- Q. Replace the cover on the crucible so that it is slightly ajar.
- R. Turn on the gas, and light the Bunsen burner.
- S. Heat the crucible for 5 minutes by slowly moving the flame back and forth across the bottom of the crucible.
- T. Place the lighted Bunsen burner under the crucible so that the flame is in contact with the

bottom of the crucible. Heat the crucible strongly for 5 minutes.

- U. Turn off the gas to extinguish flame.
- V. Remove the cover, invert it, and allow it to cool on the wire gauze. Also, remove the crucible and place it on the wire gauze to cool.
- W. When crucible and cover are cool, find the mass of the crucible, its contents, and the cover. Record this mass on line W of Data Sheet I.
- X. Discard your product according to your instructor's directions. Wash both the crucible and cover with tap water; rinse with deionized water.

Name _____

Date _____

Post-Lab Questions

1. Why should you wait until the crucible is completely cool before weighing?

3. Use the following steps to determine your percent error for the molar mass of magnesium oxide.

a. Multiply the mole ratio values of Mg and O by their respective molar masses.

$$\text{_____ Mg ratio} \bullet 24.31 \text{ g/mol} = \text{_____ g/mol}$$

$$\text{_____ O ratio} \bullet 16.00 \text{ g/mol} = \text{_____ g/mol}$$

2. How would your Mg to O ratio have been affected if:

a. the Mg was not burned completely?

Add these two amounts together in order to calculate the molar mass.

$$\text{_____ g/mol}$$

b. The accepted molar mass of magnesium oxide is 40.31 g/mol. Determine your percent error using the following equation:

b. the cover was lifted and smoke, the oxide of magnesium, escaped from the crucible?

$$\frac{(\text{accepted mass} - \text{calculated mass})}{(\text{accepted mass})} \bullet 100\%$$

$$\text{_____ \%}$$

Name _____

Date _____

Data Sheet

F. mass of crucible and cover _____ grams

H. mass of crucible, cover, and magnesium _____ grams

W. mass of crucible, cover, and magnesium oxide _____ grams

Calculations:

1. Mass of magnesium ($H - F$) _____ grams

2. Mass of oxygen ($W - H$) _____ grams

3. Moles of magnesium (molar mass of Mg is 24.31 g/mol) _____ moles

4. Moles of O (molar mass of O is 16.00 g/mol) _____ moles

5. Mole ratio of Mg (number of moles of Mg to the smaller of the two amounts of moles)

6. Mole ratio of O (number of moles of O to the smaller of the two amounts of moles)

7. Chemical formula for the oxide of magnesium

Name _____

Date _____

Pre-Lab Questions

1. Which contains more atoms: 8.00 grams of beryllium or 80.0 grams of silver? Show calculations to support your answer.

- c. Determine the number of moles of O in the compound.

_____ moles

- d. Find the mole ratio of Fe in the compound.

2. When objects made of iron, Fe, are left in the open air, they will rust. Rust, an oxide of iron is formed when the iron reacts with oxygen in the air. When 3.50 g of iron was left in the air, eventually 5.01 g of rust formed. The molar masses of iron and oxygen are 55.85 g/mole and 16.00 g/mole respectively. Find the chemical formula of this oxide of iron using the following steps:

- a. Determine the mass of O in the compound.

_____ grams

- b. Determine the number of moles of Fe in the compound.

_____ moles

- e. Find the mole ratio of O in the compound.

- f. Write the chemical formula for the rust formed, referring to Table 1 if necessary.

Chemistry of Copper and Its Quantitative Recovery

Purpose of the Experiment:

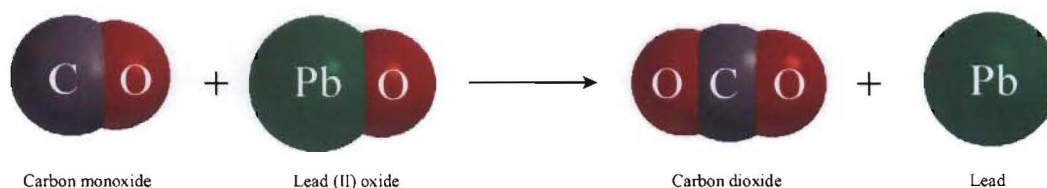
Observe chemical reactions of copper atoms and ions. Find the percent of copper recovered.

Background

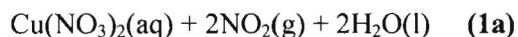
A **chemical reaction** is a chemical change in which one or more reactants (starting materials) are converted into one or more products. Chemical reactions occur all around us. They fuel and keep alive the cells of living tissues. They occur when we light a match, cook dinner, start a car, listen to a portable radio, or watch television. During a chemical reaction it may appear that a reactant is destroyed, for instance, in this experiment copper appears to be destroyed and then created again, but this is not true.

Dalton's Atomic Theory states that all matter is made up of very tiny, indivisible particles called atoms; and compounds are formed by the chemical

combination of two or more different kinds of atoms. According to the law of conservation of mass, discovered by Antoine Laurent Lavoisier, matter neither can be created nor destroyed. During a chemical reaction, the total mass of the matter at the end of the experiment is exactly the same as that at the beginning. If all matter consists of indestructible atoms, then any chemical reaction simply changes the attachments between atoms but does not destroy the atoms themselves. For example, during the reaction between carbon monoxide and lead (II) oxide, no atoms are destroyed, only rearranged. See Figure 1.

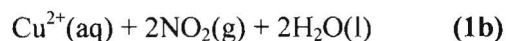
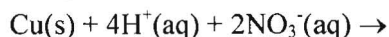


In this experiment, you will observe some of the chemistry of copper metal and its $2+$ cation, while testing your quantitative skills in recovering your starting material, the copper metal. You will begin the experiment by reacting copper metal with nitric acid in an **oxidation-reduction reaction**, where oxidation refers to the loss of electrons and reduction refers to the gain of electrons. In an oxidation-reduction reaction, electrons are transferred from one species to another.

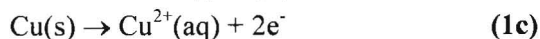


During this reaction, Cu atoms are oxidized to Cu^{2+} ions by losing two electrons. The Cu^{2+} ions are aqueous because copper (II) nitrate, $\text{Cu}(\text{NO}_3)_2$, is soluble in water. We can write a net ionic

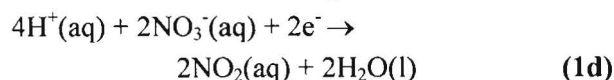
equation for this chemical reaction by canceling the spectator ions.



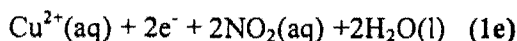
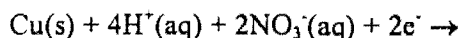
This reaction can be described as occurring in two steps, one describing the oxidation of copper and a second describing the reduction of the nitrate ion. In the oxidation step, a copper atom loses two electrons to form a copper (II) ion:



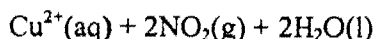
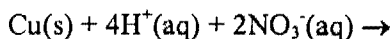
In the reduction step, two nitrate ions gain two electrons to form two nitrogen dioxide molecules:



The sum of reactions (1c) and (1d) is:

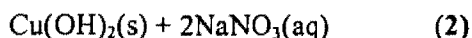
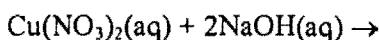


Canceling the electrons from both sides of this equation will give us our original equation (1b):

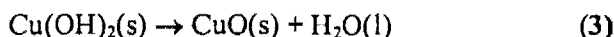


Because the copper atom transfers electrons to the nitrate ion, it is said to be the **reducing agent**. Each nitrate ion gains an electron making nitrate the **oxidizing agent**. The number of electrons lost by a reducing agent must be equal to the total number of electrons gained by an oxidizing agent. An important point to remember is that oxidizing agents and reducing agents can never be products of a reaction.

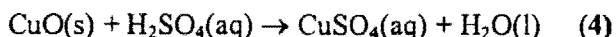
When Cu^{2+} ions react with hydroxide ions, they form copper (II) hydroxide, Cu(OH)_2 .



The resulting copper (II) hydroxide is insoluble in water, producing a colloidal suspension. Colloidal suspensions will be covered in more detail in Chapter 7. Upon heating, copper (II) hydroxide will decompose into copper (II) oxide and water.



To this point, you have taken copper from its metallic state to aqueous ions to insoluble compounds. To recover the copper in its metallic form, we will combine the copper (II) oxide, CuO , with sulfuric acid in order to produce aqueous Cu^{2+} ions.



Finally, we can recover the copper metal by adding zinc metal to the solution. The zinc atoms will react with the aqueous Cu^{2+} ions in an oxidation-reduction reaction in which electrons are transferred from the zinc metal to the Cu^{2+} ions.



Reaction (1a) has been made famous by the description attributed to Ira Remsen (ca. 1867), who read in a book that "nitric acid acts upon copper" and wanted to see what that meant.

While reading a textbook on chemistry, I came upon the statement 'nitric acid act upon copper'. I was getting tired of reading such absurd stuff and was determined to see what this meant. Copper was more or less familiar to me, for copper cents were then in use. I had seen a bottle marked 'nitric acid' on a table in the doctor's office where I was then 'doing time'! I did not know its peculiarities but I was getting on and likely to learn. The spirit of adventure was upon me. Having nitric acid and copper, I had only to learn what the words 'act upon' meant. Then, the statement, 'nitric acid acts upon copper' would be something more than mere words.

All was still. In this interest of knowledge I was even willing to sacrifice one of the few copper cents then in my possession. I put one of them on the table; opened the bottle marked 'nitric acid'; poured some of the liquid on the copper; and prepared to make an observation.

But what was this wonderful thing which I beheld? The cent was already changed, and it was no small change either. A greenish-blue liquid foamed and fumed over the cent and over the table. The air in the neighborhood of the performance became dark red. A great colored cloud arose. This was disagreeable and suffocating—how should I stop this? I tried to get rid of the objectionable mess by picking it up and throwing it out of the window, which I had meanwhile opened. I learned another fact—nitric acid not only acts upon copper but it acts upon fingers. The pain led to another unpremeditated experiment. I drew my fingers across my trousers and another fact was discovered. Nitric acid also acts upon trousers.

Taking everything into consideration, that was the most impressive experiment, and, relatively, probably the most costly experiment I have ever performed. I tell of it even now with interest. It was a revelation to me. It resulted in a desire on my part to learn more about that remarkable kind of action. Plainly, the only way to learn about it was to see its results, to experiment, to work in the laboratory.

Ira Remsen (1846-1927)

Mr. Remsen went on to found the chemistry department at Johns Hopkins University and initiated the first chemistry research center in the United States.

Procedure

- A. Obtain and weigh approximately 0.500 g of Cu wire. Add the Cu wire to a 250 mL beaker. Record the weight of the Cu wire on line A of Data Sheet I.

Caution: HNO_3 is **very** corrosive and toxic. Prevent contact with your skin, clothing, and eyes. Do not ingest.

- B. Add 4 mL of conc. HNO_3 to the beaker. **Make sure you do this under the hood.** Record your observations of the reaction on line B of Data Sheet I.

- C. After the reaction is complete, add 100 mL of deionized water. Record your observations on line C of Data Sheet I.

Caution: NaOH is corrosive and toxic. Prevent contact with your skin, clothing, and eyes. Do not ingest.

- D. Add 32 mL of 3 M NaOH to the solution in your beaker and describe the reaction on line D of Data Sheet I.

- E. Weigh 3 boiling chips. Record the mass of the boiling chips on line E of Data Sheet I. Add the boiling chips to the solution and heat on the hot plate until all the blue color is gone. While heating, continuously stir the solution with a glass rod.

- F. When the blue color has faded, remove the beaker from the hot plate and allow the CuO to settle. Describe the reaction that occurred upon heating on line F of Data Sheet I.

- G. While you are waiting for the CuO to settle, heat 200 mL of deionized water in a 400-mL beaker to 80 °C.

- H. After the CuO has settled, decant the supernatant liquid into a 600 mL beaker. (This means to pour off the liquid above the CuO solid. When pouring off the liquid, allow it to come off at a drip-by-drip pace to insure the full recovery of Cu. Try to pour out the liquid, not the solid.)

- I. Add the 200 mL of hot water to the CuO solid, stir, and allow the CuO to settle. Decant once more.

- J. Under the hood, add 13 mL of 6 M H_2SO_4 to the CuO . Record your observations on line J of Data Sheet I.

- K. Weigh out 2.0 g of Zn metal.

- L. Beneath the hood, add the zinc metal all at once and stir until the supernatant liquid is clear. Describe the reaction on line L of Data Sheet I.

- M. When gas evolution has become very slow, heat the solution gently, but do not boil, until the blue color has faded. Then, allow the solution to cool.

- N. Weigh the aspirator lid with the filter paper. Record the mass on line N of Data Sheet I.

- O. Set up the aspirator as shown by the instructor, making sure there is a good vacuum.

- P. Pour **all** of the contents in the beaker into the aspirator.

- Q. Wash the precipitated copper metal with 20 mL of deionized water.

- R. Continue to filter until the Cu appears to be dry. Weigh the aspirator lid with the filter paper, Cu, and boiling chips. Record your answer on line R of Data Sheet I.

- S. Calculate the mass of dry copper. Add the masses of the boiling chips (E) and the aspirator lid and filter paper (N). Subtract that sum from the mass of the aspirator lid, filter paper, boiling chips, and Cu (R).

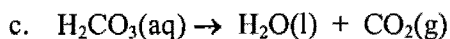
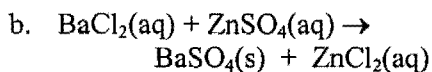
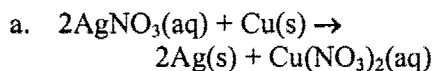
- T. Calculate the percent of copper you recovered. Divide the mass of dry copper at the end of the experiment (S) by the mass of copper at the beginning (A). Multiply by 100.

Name _____

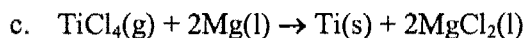
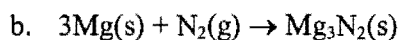
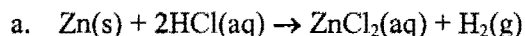
Date _____

Post-Lab Questions

1. Classify the following reactions as **combination**, **decomposition**, **double replacement**, or **single replacement**.



2. **Circle** the oxidizing agent and **underline** the reducing agent in each of the following reactions.



3. Did your results agree with Lavoisier's law of conservation of mass? If not, explain what could have gone wrong.

Name _____

Date _____

Data Sheet I

A. Mass of copper wire

A. _____ g

B. Observation after addition of concentrated nitric acid

C. Observation upon addition of water

D. Observation upon addition of NaOH

E. Mass of boiling chips

E. _____ g

F. Observation after heating

J. Observation after addition of sulfuric acid

L. Observation after addition of zinc

N. Mass of aspirator lid with filter paper

N. _____ g

R. Mass of aspirator lid, filter paper, boiling chips, and Cu

R. _____ g

S. Mass of Cu $R - (E + N)$

S. _____ g

T. Percentage of Cu recovered $(S/A) \cdot 100$

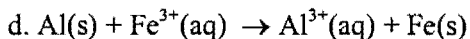
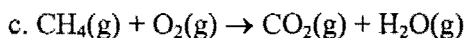
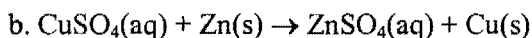
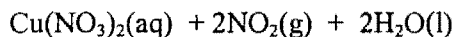
T. _____ %

Name _____

Date _____

Pre-Lab Questions

1. Circle the oxidizing agent and underline the reducing agent in each of the following reactions.



3. Classify reactions (2), (3), (4), and (5) in the lab **Background** as **combination**, **decomposition**, **double replacement**, or **single replacement**.

(2)

(3)

(4)

(5)

2. Write the net ionic equations for reactions (2), (3), (4), and (5) in the **Background** of this lab.

(2)

(3)

(4)

(5)

4. If an experiment calls for 2.50 g of aluminum at the beginning, how many grams of aluminum should be included in the product according to Lavoisier's law?